12.49 A quantity of gas is collected in a graduated tube over the mercury. The volume of the gas at 20 °C is 50.0 mL, and the level of the mercury in the tube is 100 mm above the outside mercury level. The barometer reads 750 torr. Find the volume at STP.

- Since the level of mercury inside the tube is 200 mm higher than outside, the pressure of the gas is 100 mmHg = 100 torr less than the atmosphere pressure of 750 torr.

$$\text{Volume of STP} = (50.0\ \text{mL})\left(\frac{273\ \text{K}}{293\ \text{K}}\right)\left(\frac{(750-100)\ \text{torr}}{760\ \text{torr}}\right) = 39.8\ \text{mL}$$

12.50 The respiration of a suspension of yeast cells was measured by observing the decrease in pressure of gas above the cell suspension. The apparatus was arranged so that the gas was confined to a constant volume, 16.0 cm^3, and the entire pressure change was cause by uptake of oxygen by the cells. The pressure was measured in a manometer, the fluid of which has a density of 1.034 g/cm^3. The entire apparatus was immersed in a thermostat at 37 °C. In a 30 min observation period the fluid in the open side of the manometer dropped 37 mm . Neglecting the solubility of oxygen in the yeast suspension, compute the rate of oxygen consumption by the cells in mm^3 of O_2 (STP) per hour.

- Convert the observed pressure drop to mmHg:

$$(37\ \text{mm})\left(\frac{1.034\ \text{g/cm}^3}{13.6\ \text{g/cm}^3}\right) = 2.8\ \text{mmHg in 30 min,} \quad \text{or } 5.6\ \text{mmHg}$$

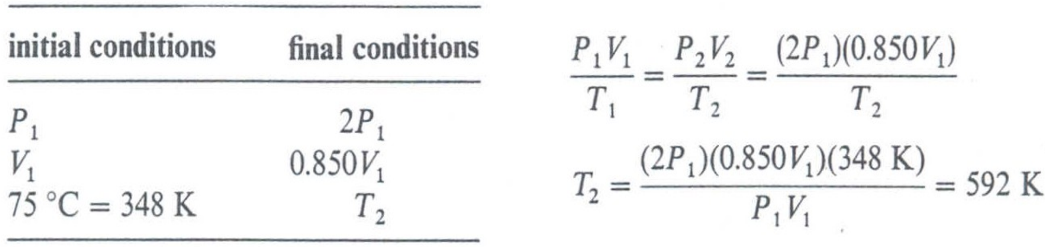
Drop per hour. With 5.6 mmHg as the oxygen partial pressure, change to STP:

$$V_2 = \frac{P_1V_1T_2}{T_1P_2} = \frac{(5.6\ \text{mmHg})(16.0\ \text{cm}^3)(273\ \text{K})}{(310\ \text{K})(760\ \text{mmHg})} = 0.10\ \text{cm}^3 = 1.0 \times 10^2\ \text{mm}^3$$

12.51 A gas at a pressure of 5.0 atm is heated from 0° to 546° and simultaneously compressed to one-third of its original volume. What will be the final pressure, in atm?

- $$P_2 = \frac{P_1V_1T_2}{T_1V_2} = \frac{(5.0\ \text{atm})(V_1)(546+273)\ \text{K}}{(273\ \text{K})(\frac{1}{3}V_1)} = 45\ \text{atm}$$

12.52 To what temperature must a neon gas sample be heated to double its pressure if the initial volume of gas at 75 °C is decreased by 15.0%?



initial conditions	final conditions
P_1	$2P_1$
V_1	$0.850V_1$
75 °C = 348 K	T_2

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{(2P_1)(0.850V_1)}{T_2}$$

$$T_2 = \frac{(2P_1)(0.850V_1)(348\text{ K})}{P_1V_1} = 592\text{ K}$$

-

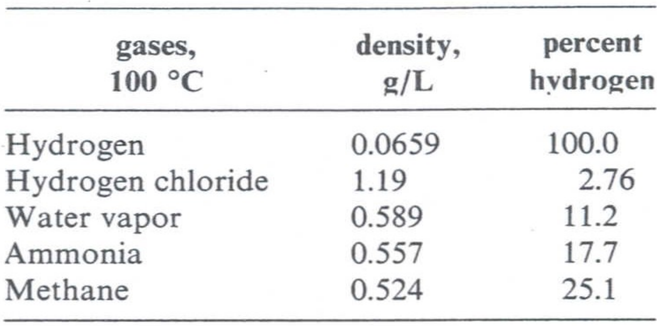
12.53 Obtain the combined gas law from Charles' law and Gay-Lussac's law.

- Charles' law says that V/T has a constant value for a given mass of gas at a given pressure. Thus V/T=f(m,P), where f is an undetermined function. Multiply the first equation by P and the second by V, to obtain two expressions, each equal to PV/T and thus equal to each other:

  Pf(m,P) = Vg(m,V)

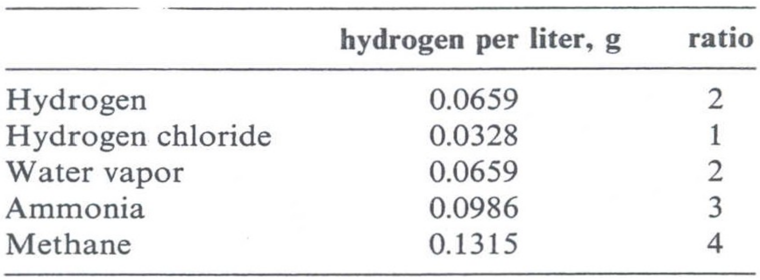
  In this last equation, when P changes, the right side does not change; therefore the left side cannot change either and must be a function of m alone. Likewise, when V changes, the left side does not change; therefore the right side is independent of V, and is a function of m alone. Hence Pf(m,P) = Vg(m,V)=h(m) and PV/T = h(m) which is the combined gas law. [the function h(m) is a constant times the number of moles of gas present.]

12.54 S.Cannizzaro (1858) was able to demonstrate that hydrogen gas consists of molecules with an even number of hydrogen atoms, based on the assumption that Avogadro's hypothesis is true. Avogadro's hypothesis states that equal volumes of gas under the same comditions of temperature and pressure contain equal numbers of molecules. Using the data from the following table, show precisely how the demonstrated this fact.



gases, 100 °C	density, g/L	percent hydrogen
Hydrogen	0.0659	100.0
Hydrogen chloride	1.19	2.76
Water vapor	0.589	11.2
Ammonia	0.557	17.7
Methane	0.524	25.1

- The number of hydrogen per L of gas, obtained by multiplying the density by the percent hydrogen, is converted to a fraction. This results are



	hydrogen per liter, g	ratio
Hydrogen	0.0659	2
Hydrogen chloride	0.0328	1
Water vapor	0.0659	2
Ammonia	0.0986	3
Methane	0.1315	4

The ratio of the number of g hydrogen per liter from one substance to another is shown in the last column. If equal volumes of all the gases contain equal numbers of molecules, and the volume of hydrogen gas has twice the number of hydrogen atoms as the same volume of hydrogen chloride (since it has twice the mass of hydrogen), then each hydrogen molecule must have twice the number of hydrogen atoms as each hydrogen chloride molecule. Thus the number of hydrogen atoms in a hydrogen molecule must be even.

12.3 MOLES OF GAS AND IDEAL GAW LAW

12.55 Which of the following, if any, contains the greatest number of N atoms? (a) 22.4 L nitrogen gas at STP (0 °C and 1 atm) (b) 500 mL of 2.00 M NH_3 (c) 1.00 mol NH_4Cl (d) 6.02 $\times 10^{23}$ molecules of NO_2.

(a) because 1 mol N_2 has 2 mol N atoms, (b) 1 mol of NH_3 has 1 mol N atoms, (c) 1 mol of NH_4Cl has 1 mol N atoms, (d) 1 mol of NO_2 has 1 mol N atoms.

12.56 Compute the weight (mass) of 6.00 L of ammonia gas, NH_3, at STP.

- $$n = \frac{6.00\ \text{L}}{22.4\ \text{L/mol (STP)}} = (0.268\ \text{mol})\left(\frac{17.0\ \text{g}}{\text{mol}}\right) = 4.55\ \text{g}$$

12.57 Determine the volume occupied by 4.0 g oxygen at STP.

- $$\text{Volume of 4.0 g } O_2 \text{ at STP} = \text{number of moles in 4.0 g } O_2 \times \text{standard molar volume}$$
$$= \left(\frac{4.0\ \text{g}}{32\ \text{g/mol}}\right)(22.4\ \text{L/mol}) = 2.8\ \text{L}$$

12.58 What volume off hydrogen will combine with 22 L chlorine to form hydrogen chloride? What volume of hydrogen chloride will be formed? Assume the same temperature and pressure for all gases.

- The balanced equation for this reaction is

$$\underset{\text{(1 molecule)}}{H_2(g)} + \underset{\text{(1 molecule)}}{Cl_2(g)} \longrightarrow \underset{\text{(2 molecules)}}{2HCl(g)}$$

The equation shows that 1 molecule H_2 reacts with 1 molecule Cl_2 to form 2 molecules HCl. But, by Avogadro's hypothesis, equal numbers of molecules of *gases* under the same conditions of temperature and pressure occupy equal volumes. Therefore the equation also indicates that 1 volume of H_2 reacts with 1 volume of Cl_2 to form 2 volumes of HCl(g). Thus 22 L of H_2 combines with 22 L of Cl_2 to form 2

of HCl.

12.59 What volume of hydrogen will unite with 16ft^3 of nitrogen to form ammonia? What volume of ammonia will be produced, all at the same temperature and pressure?

- $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

By the principles stated in Problem 12.58, 16ft^3 N_2 reacts with 3 × 16 = 48 ft^3 H_2 to form 2 x 16 = 32 ft^3 NH_3.

12.60 Calculate the volume of 0.3000 mol of gas at 60 °C and 0.821 atm.

- $V = \frac{nRT}{P} = \frac{(0.3000\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(333\ \text{K})}{0.821\ \text{atm}} = 9.99\ \text{L}$

12.61 What weight of hydrogen at STP could be contained in a vessel that holds 4.8g oxygen at STP?

- The same number of moles of H_2 as O_2 can be held in the same volume at the same temperature and pressure.

$$n(H_2) = n(O_2) = \frac{4.8\ \text{g}}{32.0\ \text{g/mol}} = 0.15\ \text{mol} \qquad (0.15\ \text{mol}\ H_2)\left(\frac{2.00\ \text{g}}{\text{mol}}\right) = 0.30\ \text{g}\ H_2$$

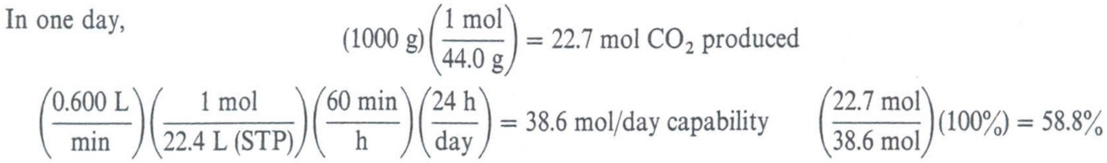
12.62 Calculate the volume of 8.40 g N_2 at t=100 °C and P=800 torr.

$$(8.40\ \text{g}\ N_2)\left(\frac{1\ \text{mol}\ N_2}{28.0\ \text{g}\ N_2}\right) = 0.300\ \text{mol}\ N_2$$

$$V = \frac{nRT}{P} = \frac{(0.300\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(373\ \text{K})}{(800/760)\ \text{atm}} = 8.73\ \text{L}$$

-

12.63 In order to economize on the oxygen supply in space ships, it has been suggested that the oxygen in exhaled CO_2 be converted to water by a reduction with hydrogen. The CO_2 output per astronaut has been estimated as 1.00 kg per 24 h day. An experimental catalytic converter reduces CO_2 at a rate 600 mL (STP) per min. What fraction of the time would such a



In one day, $(1000\ g)\left(\frac{1\ mol}{44.0\ g}\right) = 22.7\ mol\ CO_2$ produced

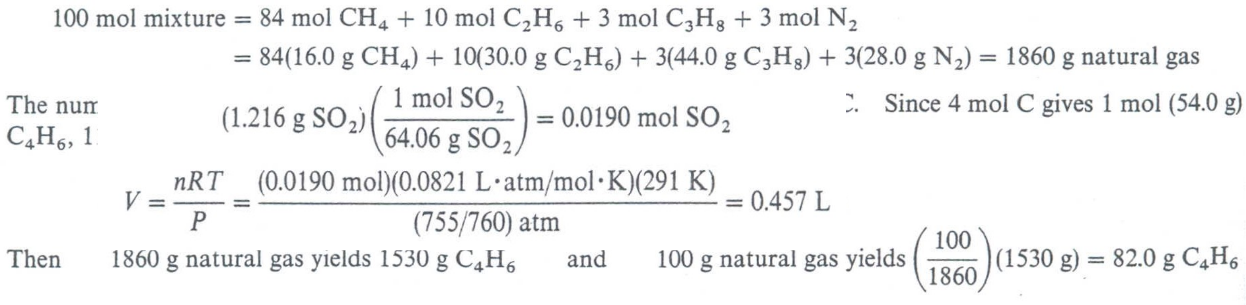
$\left(\frac{0.600\ L}{min}\right)\left(\frac{1\ mol}{22.4\ L\ (STP)}\right)\left(\frac{60\ min}{h}\right)\left(\frac{24\ h}{day}\right) = 38.6$ mol/day capability $\left(\frac{22.7\ mol}{38.6\ mol}\right)(100\%) = 58.8\%$

converter have to operate in order to keep up with CO_2 output of one astronaut?

-

12.64 A natural gas sample contains 84% (by volume) CH_4, 10% C_2H_6, 3 % C_3H_8, and 3 % N_2. If a series of catalytic reactions could be used for converting all the carbon atoms of the gas into butadiene, C_4H_6, with 100% efficiency, how much butadiene could be prepared from 100 g of the natural gas?

- Molecular weights: CH_4 = 16.0, C_2H_6 = 30.0, C_3H_8= 40, N_2=28.0, C_4H_6= 54.0. the volume percent of a gas mixture is the same as the mole percent.



100 mol mixture = 84 mol CH_4 + 10 mol C_2H_6 + 3 mol C_3H_8 + 3 mol N_2

= 84(16.0 g CH_4) + 10(30.0 g C_2H_6) + 3(44.0 g C_3H_8) + 3(28.0 g N_2) = 1860 g natural gas

The nur

C_4H_6, 1

$(1.216\ g\ SO_2)\left(\frac{1\ mol\ SO_2}{64.06\ g\ SO_2}\right) = 0.0190\ mol\ SO_2$

Since 4 mol C gives 1 mol (54.0 g)

$$V = \frac{nRT}{P} = \frac{(0.0190\ mol)(0.0821\ L\cdot atm/mol\cdot K)(291\ K)}{(755/760)\ atm} = 0.457\ L$$

Then 1860 g natural gas yields 1530 g C_4H_6 and 100 g natural gas yields $\left(\frac{100}{1860}\right)(1530\ g) = 82.0\ g\ C_4H_6$

12. 65 What volume will 1.216 g of SO_2 gas occupy at 18 °C and 755 torr?

-

12.66 Compute the volume of 14g of nitrous oxide, N_2O, at STP

- $(14\text{ g})\left(\dfrac{1\text{ mol}}{44\text{ g}}\right) = 0.32\text{ mol} \qquad V = \dfrac{nRT}{P} = \dfrac{(0.32\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(273\text{ K})}{1.00\text{ atm}} = 7.2\text{ L}$

12.67 What is the volume of 18.0 g of pure water at 1.0 atm and 4°C?

$$(18.0\text{ g})\left(\frac{1.00\text{ mL}}{1.00\text{ g}}\right) = 18.0\text{ mL}$$

-

Do not use the ideal gas law for this problem, because water at 4°C is not a gas.

12.68 Assuming the same pressure in each case, calculate the mass of hydrogen required to inflate a balloon to a certain volume V at 100°C if 3.5 g helium is required to inflate the balloon to half the volume, 0.50 V, at 25°C.

- When quantities are not known explicitly but only in relationship to other quantities, dividing one equation by another often eliminates many unknowns. For example,

$$P(H_2)V(H_2) = n(H_2)RT(H_2) \qquad \text{Dividing:} \quad \frac{P(H_2)V(H_2)}{P(\text{He})V(\text{He})} = \frac{n(H_2)RT(H_2)}{n(\text{He})RT(\text{He})}$$
$$P(\text{He})V(\text{He}) = n(\text{He})RT(\text{He})$$

But in this exercise $P(H_2) = P(\text{He})$, $\frac{1}{2}V(H_2) = V(\text{He})$, $T(H_2) = 373\text{ K}$, and $T(\text{He}) = 298\text{ K}$.

Therefore,

$$\frac{V(H_2)}{\frac{1}{2}V(H_2)} = \frac{n(H_2)(373\text{ K})}{n(\text{He})(298\text{ K})}$$

$$n(H_2) = \frac{2(298)}{373}\,n(\text{He}) = \left(\frac{2(298)}{373}\right)\left(\frac{3.5\text{ g}}{4.0\text{ g/mol}}\right) = 1.4\text{ mol }H_2 \qquad (1.4\text{ mol }H_2)(2.0\text{ g/mol}) = 2.8\text{ g }H_2$$

12.69 Liquefied natural gas (LNG) is mainly methane (MW = 16.0). A 10.0 m^3 tank is constructed to store LNG at -164°C and 1.0 atm pressure, under which conditions its density is 415 kg/m^3. Calculate the volume of a storage tank capable of holding the same mass of LNG as a gas at 20℃ and 1.0 atm pressure.

- In liquid form:

$$(10.0\ \text{m}^3)\left(\frac{415\ \text{kg}}{\text{m}^3}\right)\left(\frac{1000\ \text{mol}}{16.0\ \text{kg}}\right) = 2.59 \times 10^5\ \text{mol}$$

In gaseous form:

$$V = \frac{nRT}{P} = \frac{(2.59 \times 10^5\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(293\ \text{K})}{1.00\ \text{atm}} = 6.24 \times 10^6\ \text{L} = 6240\ \text{m}^3$$

12.70 An electronic vacuum tube was sealed off during manufacture at a pressure of 1.8 x 10^{-5} torr at 27℃. Its volume is 100 cm^3. Compute the number of gas molecules remaining in the tube.

$$n = \frac{PV}{RT} = \frac{\{[(1.8 \times 10^{-5})/(760)]\ \text{atm}\}(0.100\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(300\ \text{K})} = 9.6 \times 10^{-11}\ \text{mol}$$

$$(9.6 \times 10^{-11}\ \text{mol})(6.02 \times 10^{23}\ \text{molecules/mol}) = 5.8 \times 10^{13}\ \text{molecules}$$

-

12.71 Calculate the number of mol of gas in (a) Problem 12.17 (b) 12.18.

- (*a*) $n = \dfrac{(1.00\ \text{atm})(4.00\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(303\ \text{K})} = 0.161\ \text{mol}$

  (*b*) $n = \dfrac{(1.20\ \text{atm})(4.75\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(330\ \text{K})} = 0.210\ \text{mol}$

12.72 What volume would 25.0 g argon occupy at 90℃ and 735 torr?

- with $n = \dfrac{m}{\text{MW}}$, $V = \dfrac{mRT}{(\text{MW})P} = \dfrac{(25.0\ \text{g})(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(363\ \text{K})}{(39.9\ \text{g/mol})[(735/760)\ \text{atm}]} = 19.3\ \text{L}$

Another Method

At STP this much argon would occupy $$V_1 = \left(\frac{25.0}{39.9}\ \text{mol}\right)(22.4\ \text{L/mol}) = 14.0\ \text{L}$$

This problem can be completed by the usual procedure for converting from STP to an arbitrary set of conditions.

$$V_2 = V_1\left(\frac{T_2}{T_1}\right)\left(\frac{P_1}{P_2}\right) = (14.0\ \text{L})\left(\frac{363\ \text{K}}{273\ \text{K}}\right)\left(\frac{760\ \text{torr}}{735\ \text{torr}}\right) = 19.3\ \text{L}$$

12.73 An iron meteorite was analyzed for its isotopic argon content. The amount of ^{36}Ar was 0.200 mm^3 (STP) per kg of meteorite. If each ^{36}Ar atom had been formed by a single cosmic event, how many such events there have been per kg of meteorite?

▪ $$n = \frac{PV}{RT} = \frac{(1.00\ \text{atm})(0.200 \times 10^{-6}\ \text{L})}{(0.0821\ \text{L·atm/mol·K})(273\ \text{K})} = 8.92 \times 10^{-9}\ \text{mol}$$

$$(8.92 \times 10^{-9}\ \text{mol})\left(\frac{6.02 \times 10^{23}\ \text{atoms}}{\text{mol}}\right) = 5.4 \times 10^{15}\ \text{atoms} \qquad \text{Hence, } 5.4 \times 10^{15}\ \text{events}$$

12.74 Mercury diffusion pumps may be used in the laboratory to produce a high vacuum. Cold traps are generally placed between the pump and the system to be evacuated. These cause the condensation of mercury vapor and prevent mercury from diffusing back into the system. The minimum pressure of mercury that can exist in the system is the vapor pressure of mercury at the temperature of the cold trap. Calculate the number of mercury atoms per unit volume in a cold trap maintained at -120 °C. The vapor pressure of mercury at this temperature is 10^{-16} torr.

$$\text{Moles per liter} = \frac{n}{V} = \frac{P}{RT} = \frac{(10^{-16}/760)\ \text{atm}}{(0.0821\ \text{L·atm/K·mol})(153\ \text{K})} = 1.0 \times 10^{-20}\ \text{mol/L}$$

$$\text{Atoms per liter} = (1.0 \times 10^{-20}\ \text{mol/L})(6.0 \times 10^{23}\ \text{atoms/mol}) = 6000\ \text{atoms/L} \qquad \text{or} \qquad 6\ \text{atoms/cm}^3$$

▪

12.75 Determine the density of H_2S gas at 27 °C and 2.00 atm.

$$PV = nRT \qquad n = m/\text{MW} \qquad PV = mRT/\text{MW}$$

$$d = \frac{m}{V} = \frac{P(\text{MW})}{RT} = \frac{(2.00\ \text{atm})(34.0\ \text{g/mol})}{(0.0821\ \text{L·atm/mol·K})(300\ \text{K})} = 2.76\ \text{g/L}$$

-

12.76 Compute the approximate density of methane, CH_4, at 20 °C and 6.00 atm.

- $$d = \frac{P(\text{MW})}{RT} = \frac{(16.0\ \text{g/mol})(6.00\ \text{atm})}{(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(293\ \text{K})} = 3.99\ \text{g/L}$$

Another Method

$$\text{Density} = \frac{\text{mass of 1 mol}}{\text{volume of 1 mol}} = \frac{16.0\ \text{g}}{(22.4\ \text{L})\left(\dfrac{1.00\ \text{atm}}{6.00\ \text{atm}}\right)\left(\dfrac{293\ \text{K}}{273\ \text{K}}\right)} = 3.99\ \text{g/L}$$

12.77 A gas cylinder contains 370 g oxygen gas at 30.0 atm pressure and 25 °C. What mass of oxygen would escape if first the cylinder were heated to 75 °C and then the valve were held open until the gas pressure was 1.00 atm, the temperature being maintained at 75 °C?

- The volume is found from the initial conditions:

$$n = (370\ \text{g})\left(\frac{1\ \text{mol}}{32.0\ \text{g}}\right) = 11.6\ \text{mol}$$

$$V = \frac{nRT}{P} = \frac{(11.6\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(298\ \text{K})}{30.0\ \text{atm}} = 9.43\ \text{L}$$

The final number of moles is $n = \dfrac{PV}{RT} = \dfrac{(1.00\ \text{atm})(9.43\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(348\ \text{K})} = 0.330\ \text{mol}$

$$(0.330\ \text{mol})\left(\frac{32.0\ \text{g}}{\text{mol}}\right) = 10.6\ \text{g} \qquad (370\ \text{g initial}) - (10.6\ \text{g final}) = 359\ \text{g escaped}$$

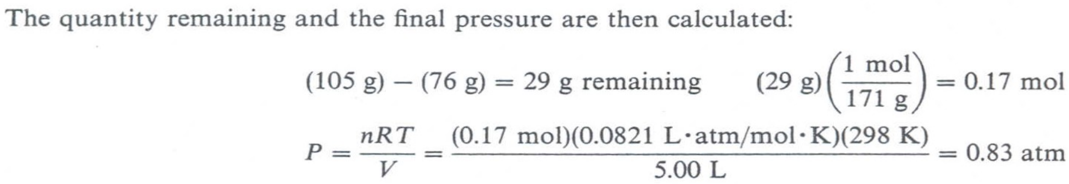
12.78 a refrigeration tank holding 5.00 L Freon gas ($C_2Cl_2F_4$) at 25 °C and 3.00 atm pressure developed a leak. When the leak was discovered and

The mass of gas initially present is calculated:

$$n = \frac{PV}{RT} = \frac{(3.00\ \text{atm})(5.00\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(298\ \text{K})} = 0.613\ \text{mol} \qquad (0.613\ \text{mol})\left(\frac{171\ \text{g}}{\text{mol}}\right) = 105\ \text{g originally present}$$

repaired, the tank had lost 76.0 g of the gas. What was the pressure of the gas remaining in the tank 25°C?

▪



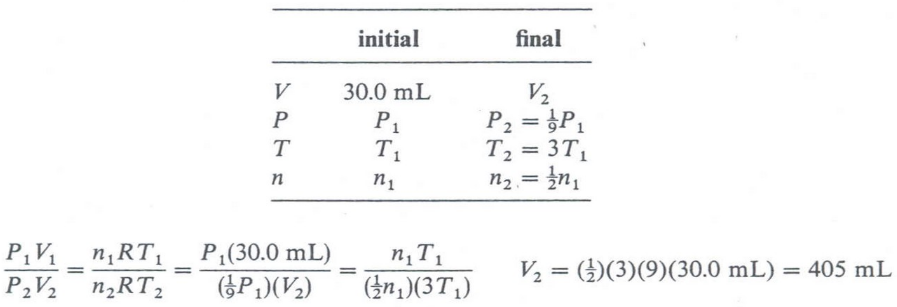
The quantity remaining and the final pressure are then calculated:

$$(105\text{ g}) - (76\text{ g}) = 29\text{ g remaining} \qquad (29\text{ g})\left(\frac{1\text{ mol}}{171\text{ g}}\right) = 0.17\text{ mol}$$

$$P = \frac{nRT}{V} = \frac{(0.17\text{ mol})(0.0821\text{ L}\cdot\text{atm/mol}\cdot\text{K})(298\text{ K})}{5.00\text{ L}} = 0.83\text{ atm}$$

12.79 A quantity of hydrogen gas occupies a volume of 30.0 mL at a certain temperature and pressure. What volume would half this mass of hydrogen occupy at triple the absolute temperature if the pressure were one-ninth that of the original gas?

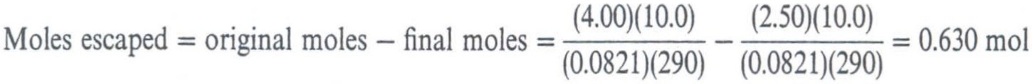
▪



	initial	final
V	30.0 mL	V_2
P	P_1	$P_2 = \frac{1}{9}P_1$
T	T_1	$T_2 = 3T_1$
n	n_1	$n_2 = \frac{1}{2}n_1$

$$\frac{P_1V_1}{P_2V_2} = \frac{n_1RT_1}{n_2RT_2} = \frac{P_1(30.0\text{ mL})}{(\frac{1}{9}P_1)(V_2)} = \frac{n_1T_1}{(\frac{1}{2}n_1)(3T_1)} \qquad V_2 = (\tfrac{1}{2})(3)(9)(30.0\text{ mL}) = 405\text{ mL}$$

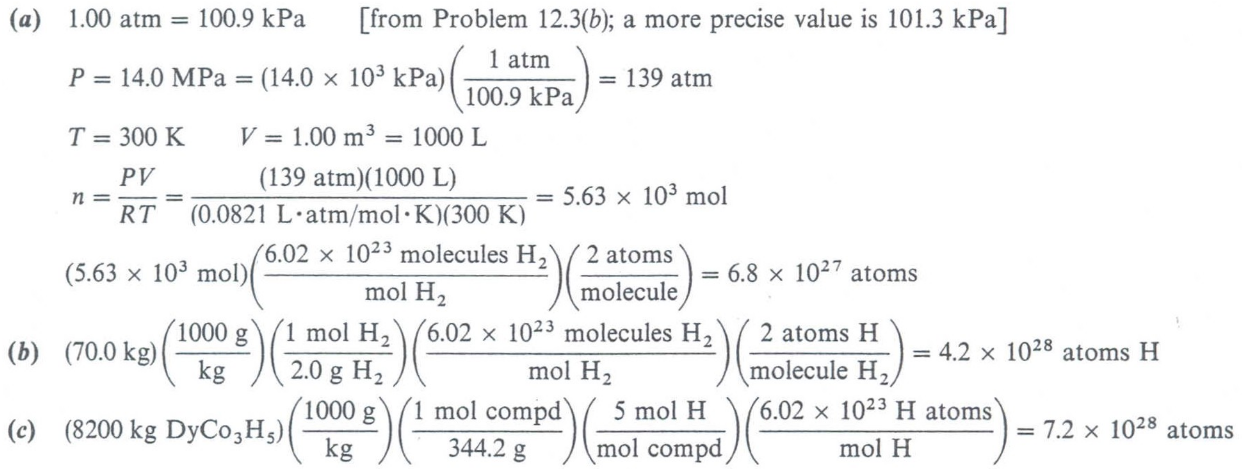
12.80 A 10.0 l cylinder of oxygen at 4.00 atm pressure and 17°C developed a leak. When the leak was repaired, 2.50 atm of oxygen remained in the cylinder, still at 17 °C. How many mol gas escaped?



$$\text{Moles escaped} = \text{original moles} - \text{final moles} = \frac{(4.00)(10.0)}{(0.0821)(290)} - \frac{(2.50)(10.0)}{(0.0821)(290)} = 0.630\text{ mol}$$

▪

12.81 One of the important attributes of hydrogen as a vehicular fuel is its compactness. Compare the numbers of the hydrogen atoms per m^3 available in (a) hydrogen gas under a pressure of 14.0 MPa at 300 K; (b) liquid hydrogen at 20 K at a density of 70.0 kg/m^3; (c) the solid compound $DyCo_3H_5$, hhaving a density of 8200 kg/m^3 at 300 K, all of whose hydrogen may be made available for combustion.



(*a*) 1.00 atm = 100.9 kPa [from Problem 12.3(*b*); a more precise value is 101.3 kPa]

$$P = 14.0\ \text{MPa} = (14.0 \times 10^3\ \text{kPa})\left(\frac{1\ \text{atm}}{100.9\ \text{kPa}}\right) = 139\ \text{atm}$$

$$T = 300\ \text{K} \qquad V = 1.00\ \text{m}^3 = 1000\ \text{L}$$

$$n = \frac{PV}{RT} = \frac{(139\ \text{atm})(1000\ \text{L})}{(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(300\ \text{K})} = 5.63 \times 10^3\ \text{mol}$$

$$(5.63 \times 10^3\ \text{mol})\left(\frac{6.02 \times 10^{23}\ \text{molecules H}_2}{\text{mol H}_2}\right)\left(\frac{2\ \text{atoms}}{\text{molecule}}\right) = 6.8 \times 10^{27}\ \text{atoms}$$

(*b*) $$(70.0\ \text{kg})\left(\frac{1000\ \text{g}}{\text{kg}}\right)\left(\frac{1\ \text{mol H}_2}{2.0\ \text{g H}_2}\right)\left(\frac{6.02 \times 10^{23}\ \text{molecules H}_2}{\text{mol H}_2}\right)\left(\frac{2\ \text{atoms H}}{\text{molecule H}_2}\right) = 4.2 \times 10^{28}\ \text{atoms H}$$

(*c*) $$(8200\ \text{kg DyCo}_3\text{H}_5)\left(\frac{1000\ \text{g}}{\text{kg}}\right)\left(\frac{1\ \text{mol compd}}{344.2\ \text{g}}\right)\left(\frac{5\ \text{mol H}}{\text{mol compd}}\right)\left(\frac{6.02 \times 10^{23}\ \text{H atoms}}{\text{mol H}}\right) = 7.2 \times 10^{28}\ \text{atoms}$$

12.82 A volume of 105 mL of pure water at 4 °C is saturated with NH_3 gas, yielding a solution of density 0.900g/mL and containing 30.0% NH_3 by weight. Find the volume of the ammonia solution resulting, and the volume of the ammonia gas at 5 °C and 775 torr, which was used to saturate the water.

$$(105\ \text{mL water})\left(\frac{1.000\ \text{g water}}{\text{mL water}}\right)\left(\frac{100\ \text{g solution}}{70.0\ \text{g water}}\right)\left(\frac{1\ \text{mL solution}}{0.900\ \text{g solution}}\right) = 167\ \text{mL}$$

$$(105\ \text{mL water})\left(\frac{1.000\ \text{g water}}{\text{mL water}}\right)\left(\frac{30.0\ \text{g NH}_3}{70.0\ \text{g water}}\right)\left(\frac{1\ \text{mol NH}_3}{17.0\ \text{g NH}_3}\right) = 2.65\ \text{mol NH}_3$$

$$V = \frac{nRT}{P} = \frac{(2.65\ \text{mol})(0.0821\ \text{L}\cdot\text{atm/mol}\cdot\text{K})(278\ \text{K})}{(775/760)\ \text{atm}} = 59.3\ \text{L}$$

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